Chemistry Enrichment Session

What we will be covering

- Topic(s):

1. Fundamental concepts encountered in GCSE and overlapping into first topic at A Level (atoms, ions, bonding, equilibria, energy).
2. GCSE Multiple choice questions covering fundamental assumed knowledge to employ from GCSE, at A Level.
3. Mathematical skills required at A Level.

How it links to the Specification:

All tasks directly link into the first module studied in A Level Chemistry, not only in a preparatory manner, but also in covering actual specification points.

Year 11 to Year 12 Transition Lesson

Stage 1— Solidify GCSE Understanding and Transition to Year 12 (take your understanding to the next level)	Stage 2— Maths Skills + Science literacy (Key definition)	Stage 3— Expand your Horizons
- Atomic structure - Structure + Bonding - Basic Organic Chemistry - Trends and Data - Oxidation and Reduction	• Quantitative Chemistry • Moles • Titrations • Balancing equations • Avogadro constant. • Ideal gas equation, including rearranging the ideal gas equation to find unknown quantities. • Students calculate percentage yields and atom economies of reactions.	Research activities Places to visit How science works in real life?

The following list comprises of fundamental chemical concepts that you have met at GCSE. Whilst we will meet these topics at A-Level again, they come up time and time again, so it is important you develop your confidence with them.

- Atomic structure
- Electron configuration
- The formation of ions and ionic charge
- Drawing ionic and covalent dot-cross diagrams and explaining the properties of different structures.
- The patterns in the periodic table including the reactions of 7 (halogens)
- Reactions between acids and bases/metals
- Calculations involving reacting masses and moles.
- Defining relative molecular and atomic mass and calculating the Mr/Ar for different substances
- Defining oxidation and reduction in terms of electrons and writing balanced ionic equations
- Precipitation reactions
- Equilibrium
- Energy changes (exothermic/ endothermic reactions) and rates of reactions
- Organic chemistry—Naming compounds and identifying functional groups

Addressing key concepts to solidify GCSE understanding and how it links to Year 12

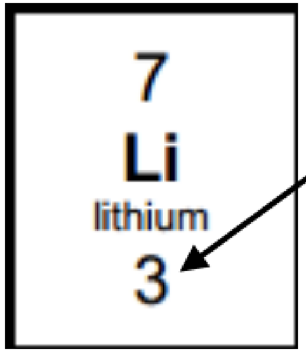
Pre-Knowledge Topics

Chemistry topic 1 – Electronic structure, how electrons are arranged around the nucleus

A periodic table can give you the proton / atomic number of an element, this also tells you how many electrons are in the ***atom***.

You will have used the rule of electrons shell filling, where:

The first shell holds up to 2 electrons, the second up to 8, the third up to 8 and the fourth up to 18 (or you may have been told 8).



Atomic number =3, electrons = 3, arrangement 2 in the first shell and 1 in the second or

Li = 2,1

At **A level** you will learn that the electron structure is more complex than this, and can be used to explain a lot of the chemical properties of elements.

The 'shells' can be broken down into 'orbitals', which are given letters:'s' orbitals, 'p' orbitals and 'd' orbitals.

- ***The first period***

Hydrogen has its only electron in the 1s orbital - $\mathbf{1s^1}$, and at helium the first level is completely full - $\mathbf{1s^2}$.

The second period

Now we need to start filling the second level, and hence start the second period.

Lithium's electron goes into the 2s orbital because that has a lower energy than the 2p orbitals. Lithium has an electronic structure of $1s^2\mathbf{2s^1}$. Beryllium adds a second electron to this same level - $1s^2\mathbf{2s^2}$. Now the 2p levels start to fill. These levels all have the same energy, and so the electrons go in singly at first.

B	$1s^22s^2\mathbf{2p_x^1}$
C	$1s^22s^22p_x^1\mathbf{2p_y^1}$
N	$1s^22s^22p_x^12p_y^1\mathbf{2p_z^1}$

- You can see that it is going to get progressively tedious to write the full electronic structures of atoms as the number of electrons increases. There are two ways around this, and you must be familiar with both.
- ***Shortcut 1:*** All the various p electrons can be lumped together. For example, fluorine could be written as $1s^2 2s^2 2p^5$, and neon as $1s^2 2s^2 2p^6$

Shortcut 2: You can lump *all* the inner electrons together using, for example, the symbol [Ne]. In this context, [Ne] means *the electronic structure of neon* - in other words: $1s^2 2s^2 2p_x^2 2p_y^2 2p_z^2$

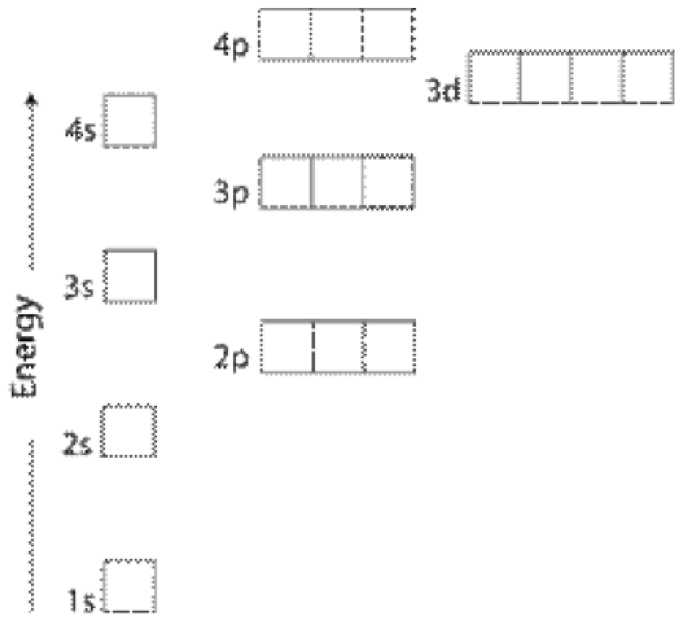
You wouldn't do this with helium because it takes longer to write [He] than it does $1s^2$.
On this basis the structure of chlorine would be written $[Ne]3s^2 3p_x^2 3p_y^2 3p_z^1$.

For example:

		short version
Mg	$1s^2 2s^2 2p^6 3s^2$	$[Ne]3s^2$
S	$1s^2 2s^2 2p^6 3s^2 3p_x^2 3p_y^1 3p_z^1$	$[Ne]3s^2 3p_x^2 3p_y^1 3p_z^1$
Ar	$1s^2 2s^2 2p^6 3s^2 3p_x^2 3p_y^2 3p_z^2$	$[Ne]3s^2 3p_x^2 3p_y^2 3p_z^2$

The arrangement of electrons within an atom is called the electronic configuration and the electrons are filled up according to the energy of the levels as: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f.

Filling up of the electrons and the resulting electronic configuration of an element is governed by some rules which are pivotal to the understanding of the chemical processes.



Its observed that everything looks correct until the 3p subshell is passed. It is expected that the 3p subshell will be followed by 3d, but it is really the 4s which follows.

Although their energies are close, 4s is lower in energy because of electrons in this orbital being able to move close to the nucleus due to the symmetrical spherical shape of the s orbital.

The electrons in the 3d orbitals are not able do this and hence are higher in energy.

Summary

Writing the electronic structure of an element from hydrogen to krypton

- Use the Periodic Table to find the atomic number, and hence number of electrons.
- Fill up orbitals in the order 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p - until you run out of electrons.
- The 3d is the awkward one - remember that specially. Fill p and d orbitals singly as far as possible before pairing electrons up.

Now that you are familiar with s, p and d orbitals try these problems, write your answer in the format:

$1s^2$, $2s^2$, $2p^6$ etc.

Q1.1 Write out the electron configuration of:

a) Ca b) Al c) S d) Cl e) Ar f) Fe g) V h) Ni i) Cu j) Zn k) As

Q1.2 Extension question, can you write out the electron arrangement of the following ***ions***:

a) K^+ b) O^{2-} c) Zn^{2+} d) V^{5+} e) Co^{2+}

Pre-Knowledge Topics Answers to problems

Q1.1a) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ b) $1s^2 2s^2 2p^6 3s^2 3p^1$ c) $1s^2 2s^2 2p^6 3s^2 3p^4$ d) $1s^2 2s^2 2p^6 3s^2 3p^5$

e) $1s^2 2s^2 2p^6 3s^2 3p^6$ f) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$ g) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$

h) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$ i) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$ j) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$

k) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$

Q1.2a) $1s^2 2s^2 2p^6 3s^2 3p^6$ b) $1s^2 2s^2 2p^6 3s^2 3p^6$ c) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$

d) $1s^2 2s^2 2p^6 3s^2 3p^6$ e) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^7$

Chemistry topic 2 – Oxidation and reduction

At GCSE you know that oxidation is adding oxygen to an atom or molecule and that reduction is removing oxygen, or that oxidation is removing hydrogen and reduction is adding hydrogen. You may have also learned that oxidation is removing electrons and reduction is adding electrons.

At A level we use the idea of oxidation number a lot!

You know that the metals in group 1 react to form ions that are +1, i.e. Na+ and that group 7, the halogens, form -1 ions, i.e. Br-. We say that sodium, when it has reacted has an oxidation number of +1 and that bromide has an oxidation number of -1. All atoms that are involved in a reaction can be given an oxidation number.

An element, Na or O2 is always given an oxidation state of zero (0), any element that has reacted has an oxidation state of + or -. As removing electrons is reduction, if, in a reaction the element becomes more negative it has been reduced, if it becomes more positive it has been oxidised.

Elements that you expect to have a specific oxidation state have different states, so for example you would expect chlorine to be -1, it can have many oxidation states: NaClO, in this compound it has an oxidation state of +1

There are a few simple rules to remember:

Metals have a + oxidation state when they react.

Oxygen is 'king' it always has an oxidation state of -2

Hydrogen has an oxidation state of +1 (except metal hydrides)

The charges in a molecule must cancel.

Examples: Sodium nitrate, $NaNO_3$

Na +1 3x O^{2-}

+1 -6

To cancel: N = +5

sulfate ion, SO_4^{2-}

4xO^{2-} and 2- charges 'showing'

-8 -2

S = +6

Q2.1 Work out the oxidation state of the underlined atom in the following:

a) $Mg\underline{C}O_3$ b) $\underline{S}O_3$ c) $Na\underline{Cl}O_3$ d) $\underline{Mn}O_2$ e) $\underline{Fe}_2O_3$ f) $\underline{V}_2O_5$

g) $K\underline{Mn}O_4$ h) $\underline{Cr}_2O_7^{2-}$ i) $\underline{Cl}_2O_4$

Q2.1 a) +4 b) +6 c) +5 d) +4 e) +3 f) +5 g) +7 h) +6 i) +4

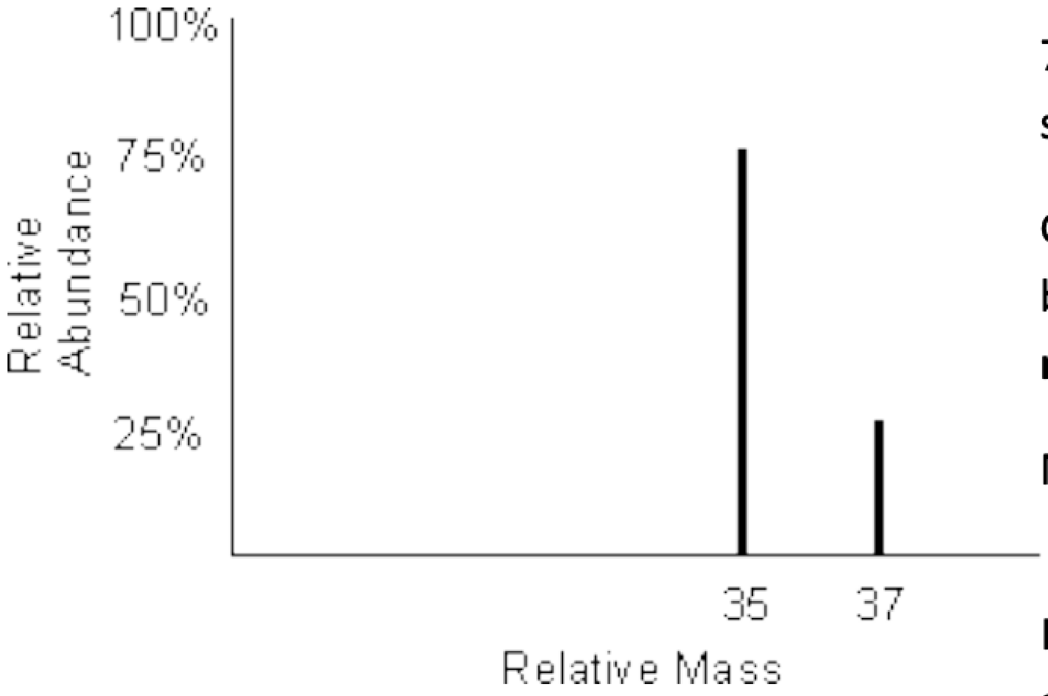
Chemistry topic 3 – Isotopes and mass

- You will remember that an isotopes are elements that have differing numbers of neutrons. Hydrogen has 3 isotopes H_1^1 H_1^2 H_1^3
- Isotopes occur naturally, so in a sample of an element you will have a mixture of these isotopes. We can accurately measure the amount of an isotope using a mass spectrometer.
- You will need to understand what a mass spectrometer is and how it works at A level.

Q3.1 What must happen to the atoms before they are accelerated in the mass spectrometer?

Q3.2 Explain why the different isotopes travel at different speeds in a mass spectrometer.

A mass spectrum for the element chlorine will give a spectrum like this:



75% of the sample consist of chlorine-35, and 25% of the sample is chlorine-37.

Given a sample of naturally occurring chlorine ¾ of it will be Cl-35 and ¼ of it is Cl-37. We can calculate what the **mean** mass of the sample will be:

$$\text{Mean mass} = \frac{75}{100} \times 35 + \frac{25}{100} \times 37 = 35.5$$

If you look at a periodic table this is why chlorine has an atomic mass of 35.5.

Q3.1 They must be ionised / turned into ions

Q3.2 The ions are all given the same amount of kinetic energy, as $KE = \frac{1}{2} mv^2$ the lighter ions will have greater speed / heavier ions will have less speed.

An A level periodic table has the masses of elements recorded much more accurately than at GCSE. Most elements have isotopes, and these have been recorded using mass spectrometers.

GCSE

11 **B** boron 5	12 **C** carbon 6	14 **N** nitrogen 7	16 **O** oxygen 8	19 **F** fluorine 9
27 **Al** aluminium 13	28 **Si** silicon 14	31 **P** phosphorus 15	32 **S** sulfur 16	35.5 **Cl** chlorine 17

A level

10.8 **B** 5 boron	12.0 **C** 6 carbon	14.0 **N** 7 nitrogen	16.0 **O** 8 oxygen	19.0 **F** 9 fluorine
27.0 **Al** 13 aluminium	28.1 **Si** 14 silicon	31.0 **P** 15 phosphorus	32.1 **S** 16 sulphur	35.5 **Cl** 17 chlorine

Given the percentage of each isotope you can calculate the mean mass which is the accurate atomic mass for that element.

Q3.3 Use the percentages of each isotope to calculate the accurate atomic mass of the following elements.

a) Antimony has 2 isotopes: Sb-121 57.25% and Sb-123 42.75%
b) Gallium has 2 isotopes: Ga-69 60.2% and Ga-71 39.8%
c) Silver has 2 isotopes: Ag-107 51.35% and Ag-109 48.65%
d) Thallium has 2 isotopes: Tl-203 29.5% and Tl-205 70.5%
e) Strontium has **4** isotopes: Sr-84 0.56%, Sr-86 9.86%, Sr-87 7.02% and Sr-88 82.56%

- Q3.3
- a) 121.855
- b) 67.796
- c) 107.973
- d) 204.41
- e) 87.710 / 87.7102

Chemistry topic 5 – Measuring chemicals – the mole

- From this point on you need to be using an A level periodic table, not a GCSE one you can view one here: http://bit.ly/pixlpertab
- Now that we have our chemical equations balanced, we need to be able to use them in order to work out masses of chemicals we need or we can produce.

The mole

- The ***mole*** is the chemists equivalent of a dozen, atoms are so small that we cannot count them out individually, we weigh out chemicals.

For example: magnesium + sulfur → magnesium sulfide

$Mg + S \rightarrow MgS$

We can see that one atom of magnesium will react with one atom of sulfur, if we had to weigh out the atoms we need to know how heavy each atom is.

From the periodic table: Mg = 24.3 and S = 32.1

If I weigh out exactly 24.3g of magnesium this will be 1 mole of magnesium, if we counted how many atoms were present in this mass it would be a huge number (6.02×10^{23}!!!!), if I weigh out 32.1g of sulfur then I would have 1 mole of sulfur atoms.

So 24.3g of Mg will react precisely with 32.1g of sulfur, and will make 56.4g of magnesium sulfide.

The Mole & Avogadro's Constant

- The mole is the unit representing the amount of atoms, ions, or molecules
- One mole is the amount of a substance that contains **6.02×10^{23}** particles (Atoms, Molecules or Formulae) of a substance (6.02×10^{23} is known as the **Avogadro Number)**

Examples

- 1 mole of Sodium (Na) contains 6.02×10^{23} **Atoms** of Sodium
- 1 mole of Hydrogen (H_2) contains 6.02×10^{23} **Molecules** of Hydrogen
- 1 mole of Sodium Chloride (NaCl) contains 6.02×10^{23} **Formula units** of Sodium Chloride

Linking the mole and the atomic mass

- One mole of any element is equal to the relative atomic mass of that element in grams
- For example one mole of carbon, that is if you had 6.02×10^{23} atoms of carbon in your hand, it would have a mass of 12g
- So one mole of helium atoms would have a mass of 4g, lithium 7g etc
- For a compound we add up the relative atomic masses
- So one mole of water would have a mass of $2 \times 1 + 16 = 18$g
- Hydrogen which has an atomic mass of 1 is therefore equal to $^{1}/_{12}$ the mass of a ^{12}C atom
- So one carbon atom has the same mass as 12 hydrogen atoms

3.1.2.2 - Moles and the Avogadro Constant

The mole is a **unit of measurement** for substances. It always contains the **same number of particles**.

$$L = 6.022 \times 10^{23} \text{ particles}$$

This number is the **Avogadro Constant** (L) and allows the number of particles present in a sample of a substance with known mass to be found:

$$\text{Number of particles} = nL$$

(n = moles) (L = Avogadro constant)

The mole is a **very important unit of measurement** in many calculations:

$$\text{Moles} = \frac{\text{mass}}{\text{Mr}} = \frac{\text{concentration x volume}}{1000}$$

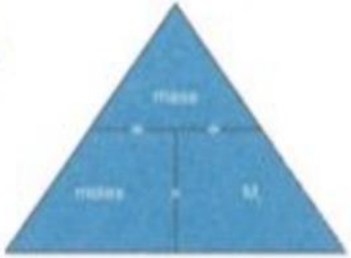
(where concentration is in $moldm^{-3}$)

Chemistry Topic 5 - Formulae and Equations

MOLE ⇒ The amount of substance IN GRAMS that has the same number of particles as there are atoms in 12 grams of carbon-12.

FOR SOLIDS...

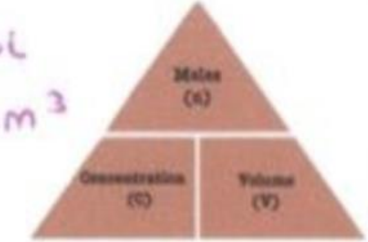
$\frac{\text{mass(g)}}{M_r}$



FOR LIQUIDS...

conc × vol

$mol\,dm^{-1} \times dm^3$

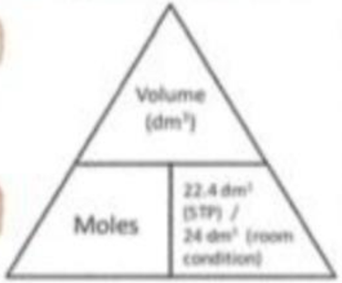


FOR GASES volume needs to be in dm^3!!!

$\frac{vol}{24}$ (25°c)

OR

$\frac{vol}{22.4}$ (STP)



AVOGADRO'S CONSTANT...

$mol \times 6.02 \times 10^{23} = n^o$ of particles

e.g. 1 mole of copper will have 6.02×10^{23} atoms.

1 mole of CO_2 will have 6.02×10^{23} molecules.

(always the same)

EMPIRICAL FORMULA

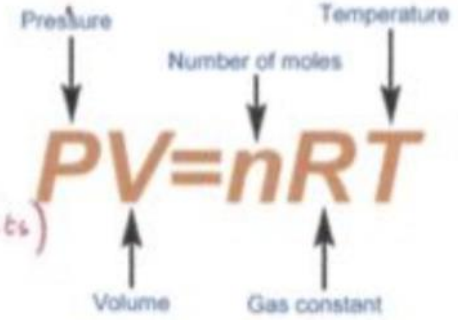
Simplest ratio of atoms in compound

① Divide each mass (or %) by

IDEAL GAS EQUATION

Same as the above triangle but can give gas volume for any temperature and pressure ☺

moles = $\frac{P \times V}{R \times T}$ (remember units)



Pressure ⇒ Pa (KPa → Pa = ×1000)

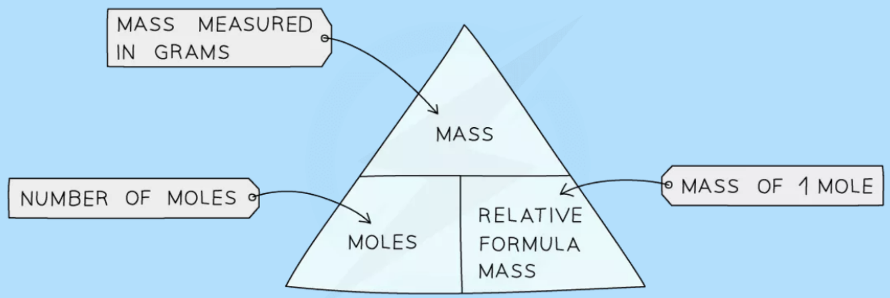
Volume ⇒ M^3 ($cm^3 \rightarrow m^3$ = ÷ 1 million)

Gas Constant ⇒ 8.31 ($J.mol^{-1}.K^{-1}$)

Temperature ⇒ Kelvin (°c → K = +273)

Calculating Reacting Masses, Solutions & Concentrations of Solutions in g/dm3 & mol/dm3

Calculating percentage composition, moles, mass and relative formula mass



1. Calculating Moles

Equation:

Amount in Moles = Mass of Substance in grams ÷ M_r (or A_r)

Example:

SUBSTANCE	MASS	MR	AMOUNT
NaOH	80 g	40	(80 ÷ 40) = 2 moles
$CaCO_3$	25 g	100	(25 ÷ 100) = 0.25 moles
H_2SO_4	4.9 g	98	(4.9 ÷ 98) = 0.05 moles
H_2O	108 g	18	(108 ÷ 18) = 6 moles
$CuSO_4.5H_2O$	75 g	250	(75 ÷ 250) = 0.3 moles

2. Calculating Mass

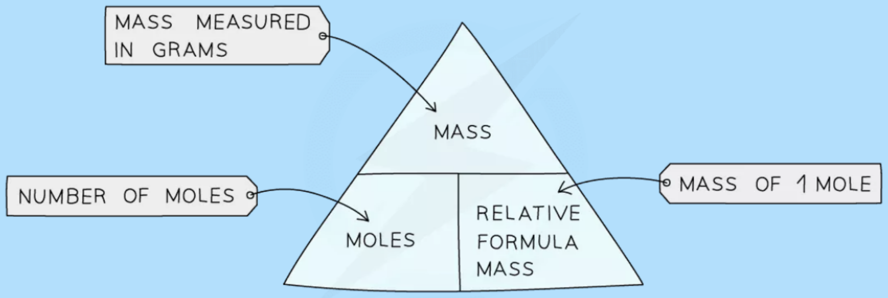
Equation:

Mass of substance (grams) = Moles x M_r (or A_r)

Example:

SUBSTANCE	AMOUNT	MR	MASS
H_2O	0.5 moles	18	(0.5 x 18) = 9 g
NaCl	3 moles	58.5	(3 x 58.5) = 175.5 g
K_2CO_3	0.2 moles	138	(0.2 x 138) = 27.6 g
$(NH_4)2SO_4$	2.5 moles	132	(2.5 x 132) = 330 g
$MgSO_4.7H_2O$	0.25 moles	246	(0.25 x 246) = 61.5 g

Calculating percentage composition, moles, mass and relative formula mass



3. Calculating Relative Formula Mass

Equation:

M_r (or A_r) = Mass of Substance in Grams ÷ Moles

Example:

10 moles of Carbon Dioxide has a Mass of 440 g. What is the Relative Formula Mass of Carbon Dioxide?

Relative Formula Mass = Mass ÷ Number of Moles

Relative Formula Mass = 440 ÷ 10 = 44

Relative Formula Mass of Carbon Dioxide = 44

4. Calculating Percentage Composition

- The percentage composition is found by calculating the percentage by mass of each element in a compound

Example:

Calculate the percentage of oxygen in CO_2

Step 1 – Calculate the molar mass of the compound

Molar mass CO_2 = (2 x 16) + 12 = 44

Step 2 – Add the atomic masses of the element required as in the question (oxygen)

16 + 16 = 32

Step 3 – Calculate the percentage

% of oxygen in CO_2 = 32/44 x 100 = 72.7%

Q6.1 Answer the following questions on moles.

a) How many moles of phosphorus pentoxide (P_4O_{10}) are in 85.2g?
b) How many moles of potassium in 73.56g of potassium chlorate (V) ($KClO_3$)?

c) How many moles of water are in 249.6g of hydrated copper sulfate(VI) ($CuSO_4.5H_2O$)? For this one, you need to be aware the dot followed by $5H_2O$ means that the molecule comes with 5 water molecules so these have to be counted in as part of the molecules mass.
d) What is the mass of 0.125 moles of tin sulfate ($SnSO_4$)?
e) If I have 2.4g of magnesium, how many g of oxygen(O_2) will I need to react completely with the magnesium? $2Mg + O_2 \rightarrow MgO$

- a) 85.2/284 = 0.3 moles
- b) 73.56/122.6 = 0.6 moles
- c) 249.5/249.5 = 1.0 moles
- d) 0.125 x 212.8 = 26.6g
- e) 2Mg : 2O or 1:1 ratio 2.4g of Mg = 0.1moles so we need 0.1 moles of oxygen (O2): 0.1 x 32 = 3.2g

2. Calculating Concentration

Equation:

```
Concentration (mol / dm³)  =  Amount of substance (mol) ÷ Volume of solution (dm³)
```

Example:

25.0 cm^3 of 0.050 mol / dm^3 sodium carbonate was completely neutralised by 20.00 cm^3 of dilute hydrochloric acid. Calculate the concentration in mol / dm^3 of the hydrochloric acid.

Step 1 – Calculate the amount, in moles, of sodium carbonate reacted by rearranging the equation for amount of substance (mol) and dividing by 1000 to convert cm^3 to dm^3

Amount of Na_2CO_3 = (25.0 x 0.050) ÷ 1000 = 0.00125 mol

Step 2 – Calculate the amount, in moles, of hydrochloric acid reacted

1 mol of Na_2CO_3 reacts with 2 mol of HCl, so the Molar Ratio is 1 : 2

Therefore 0.00125 moles of Na_2CO_3 react with 0.00250 moles of HCl

Step 3 – Calculate the concentration, in mol / dm^3 of the Hydrochloric Acid

1 dm^3 = 1000 cm^3

Volume of HCl = 20 ÷ 1000 = 0.0200 dm^3

Concentration HCl (mol / dm^3) = 0.00250 ÷ 0.0200 = 0.125

Concentration of Hydrochloric Acid = 0.125 mol / dm^3

3. Calculating Volume

Equation:

Volume (dm^3) = Amount of substance (mol) ÷ Concentration (mol / dm^3)

Example:

Calculate the volume of hydrochloric acid of concentration 1.0 mol / dm^3 that is required to react completely with 2.5g of calcium carbonate.

Step 1 – Calculate the amount, in moles, of calcium carbonate that reacts

M_r of $CaCO_3$ is 100

Amount of $CaCO_3$ = (2.5 ÷ 100) = 0.025 mol

Step 2 – Calculate the moles of hydrochloric acid required

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$

1 mol of $CaCO_3$ requires 2 mol of HCl

So 0.025 mol of $CaCO_3$ Requires 0.05 mol of HCl

Step 3 – Calculate the volume of HCl Required

Volume = (Amount of Substance(mol) ÷ Concentration (mol / dm^3)

= 0.05 ÷ 1.0

= 0.05 dm^3 (the moles cancel out above and below the line)

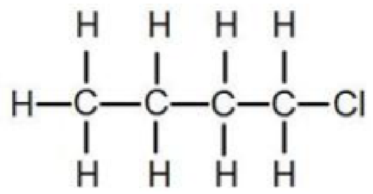
Volume of Hydrochloric Acid = 0.05 dm^3

Chemistry topic – Organic chemistry – functional groups

- At GCSE you would have come across hydrocarbons such as alkanes (ethane etc) and alkenes (ethene etc). You may have come across molecules such as alcohols and carboxylic acids.
- At A level you will learn about a wide range of molecules that have had atoms added to the carbon chain. These are called functional groups, they give the molecule certain physical and chemical properties that can make them incredibly useful to us.
- Here you are going to meet a selection of the functional groups, learn a little about their properties and how we give them logical names.
- You will find a menu for organic compounds here:
- http://bit.ly/pixlchem13

Using the two links see if you can answer the following questions:

Q9.1 Halogenoalkanes



What is the name of this halogenoalkane?

How could you make it from butan-1-ol?

Q9.2 Alcohols

How could you make ethanol from ethene?

How does ethanol react with sodium, in what ways is this a) similar to the reaction with water, b) different to the reaction with water?

Q9.3 Aldehydes and ketones

Draw the structures of a) propanal b) propanone

How are these two functional groups different?

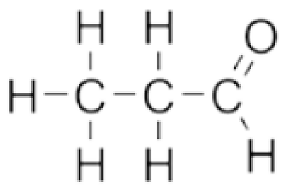
9.1 1-chlorobutane

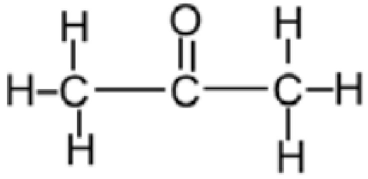
Add butan-1-ol to concentrated HCl and shake

9.2 react ethene with hydrogen gas at high temperature and pressure with a nickel catalyst

The reaction is similar in that it releases hydrogen but different as it proceeds much slower than in water

9.3 propanal propanone





The carbon atom joined to oxygen in propanal has a hydrogen attached to it, it does not in propanone.

==

Research activity

Use your online searching abilities to see if you can find out as much about the topic as you can. Remember it you are a prospective A level chemist, you should aim to push your knowledge.

- Task 1: The chemistry of fireworks What are the component parts of fireworks? What chemical compounds cause fireworks to explode? What chemical compounds are responsible for the colour of fireworks?
- Task 2: Why is copper sulfate blue? Copper compounds like many of the transition metal compounds have got vivid and distinctive colours – but why?
- Task 3: Aspirin What was the history of the discovery of aspirin, how do we manufacture aspirin in a modern chemical process?
- Task 4: The hole in the ozone layer Why did we get a hole in the ozone layer? What chemicals were responsible for it? Why were we producing so many of these chemicals? What is the chemistry behind the ozone destruction?

All data is given on this paper, you will not need a periodic table

Answer all questions.

1. Here is part of a periodic table, use it to answer the following questions

10.8 $_5$B boron	12.0 $_6$C carbon	14.0 $_7$N nitrogen	16.0 $_8$O oxygen	19.0 $_9$F fluorine	20.2 $_{10}$Ne neon
27.0 $_{13}$Al aluminium	28.1 $_{14}$Si silicon	31.0 $_{15}$P phosphorus	32.1 $_{16}$S sulphur	35.5 $_{17}$Cl chlorine	39.9 $_{18}$Ar argon

a. Which is the correct electron configuration for a nitrogen atom, circle the correct answer [1]

$1s^2 2p^5$ $1s^1 2p^6$ $1s^2 2s^2 2p^3$ $1s^2 2s^5$ $1s^2 2s^2 2p^6 3s^2 3p^2$

b. Which is the correct electron configuration for a chlorine atom, circle the correct answer [1]

$1s^2 2s^8 2p^7$ $1s^2 2s^2 2p^8 2d^5$ $1s^2 2s^2 2p^6 3d^7$ $1s^2 2s^2 2p^6 3p^7$ $1s^2 2s^2 2p^6 3s^2 3p^5$

c. Which is the correct electron configuration for an aluminium **ion**, Al^{3+}? Circle the correct answer [1]

$1s^2 2s^2 2p^6$ $1s^2 2s^2 2p^6 3s^2 3p^3$ $1s^2 2s^2 2p^6 3s^2$ $1s^2 2s^2 2p^6 2d^1$

2. Draw a dot and cross diagram to show the bonding in a molecule of water, H_2O. [2]
Atomic numbers: H =1, O =8

4. A mass spectrometer was used to analyse a sample of chlorine; the results of the analysis are as follows:

isotope mass	% of sample
Cl-35	75.53
Cl-37	24.47

Calculate the accurate atomic mass of chlorine. Give your answer to **3 decimal places**. [3]

mass: ______________________

5. Give the oxidation state of the underlined atom in the following chemicals.
Useful information: H = +1, K = +1, Na = +1, Mg = +2, O = -2, Cl = -1 [7]

a) $\underline{C}O_2$ b) $\underline{S}O_3$ c) $H_2\underline{S}O_4$ d) $\underline{Al}Cl_3$

e) $\underline{Cr}_2O_3$ f) $Na\underline{N}O_3$ g) $\underline{V}Cl_4$

6. Balance the following chemical equations:

a) C_3H_8 + ___ O_2 $\rightarrow$ ___ CO_2 + ___ H_2O [3]

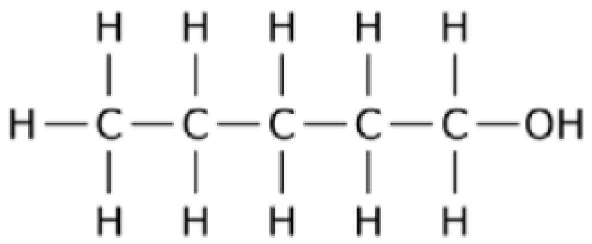
b) ___ HCl + $Mg(OH)_2$ $\rightarrow$ $MgCl_2$ + ___ H_2O [2]

c) Na_2CO_3 + ___ HCl $\rightarrow$ ___ NaCl + ___ H_2O + CO_2 [3]

7. Calculate the relative formula masses of the following:
Atomic masses: H = 1, O = 16, S = 32.1, C = 12, Ca = 40.1, Na = 23, Cl = 35.5, Zn = 65.4

a) $CaCl_2$ b) H_2CO_3 c) Na_2SO_4 d) C_3H_7OH e) $Zn(NO_3)_2$ [5]

8. A student carried out a reaction with this molecule:



a. What is the name of this molecule? ____________________ [2]